Chapter 5: Electrons in Atoms

5.1 Light and Quantized Energy Ideas about Matter

I. The Atom and unanswered questions

A. In early 1900s, scientists began to unravel the puzzle of chemical behavior

1. Scientists noticed certain elements emitted visible light and that an element's chemical behavior is related to the electrons in its atoms

II The Wave Nature of Light

A. Visible light is a type of electromagnetic radiation

1. ex: microwaves, X-rays, radio waves

B. Characteristics of Waves

1. Wavelength - measured from crest to crest, shortest distance between equal pts. on crests

2. Frequency - # of waves that pass a given pt. per second

3. Amplitude - wave's height from origin to crest

a. Wavelength + frequency are not effected by amplitude

C. Electromagnetic Wave relationship = $c = \lambda \nu$

D. Electromagnetic Spectrum (EM spectrum) - includes all forms of EM radiation

→ See Example Problem 1 p.273

c = speed of light = 3.00×10^8 m/s
λ = wavelength
ν = freq.

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III Particle Nature of Light

A. matter can gain or lose energy only in small, specific amounts called quanta: a quantum is a standard used to measure this

B. Max Planck - proposed energy emitted by hot objects was quantized

1. Energy of a quantum → $E_{quantum} = h\nu$

a. h = Planck's constant → 6.626×10^{-34} J·s

b. Joules = SI unit for energy

C. Photoelectric Effect - photoelectrons are emitted from a metal's surface when a certain frequency shines on its surface

D. Light's dual nature

1. A beam of light can be thought of as a beam of bundles of energy called photons

a. Energy of photon → $E_{photon} = h\nu$

b. even small numbers of photons with energy above the threshold ejects photoelectrons

IV Atomic Emission Spectra

A. Each element has a unique atomic emission spectra that can be used to identify it

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5.1 Review

Grade: 11th
Subject: Chem.
Date:

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1 Put in order of increasing wavelength for the following types of electromagnetic radiation. Ultraviolet (UV) rays, microwaves, radio waves, X-rays.

A x-rays, radio waves, UV rays, microwaves
B Radio waves, microwaves, UV rays, X-rays

C X-rays, UV rays, microwaves, radio waves

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2 The number of waves that pass a given point per second is defined as what?

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3 The shortest distance between equivalent points on a continuous wave is defined as what?

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4 What is the wavelength of electromagnetic radiation with a frequency of 5.00 x 10^12 Hz ?

$\frac{3.0\times 8}{5.0\times 10^{12}} = \lambda \cdot \frac{5.0\times 10^{12}}{5.0\times 10^{12}} \rightarrow \lambda = 6.0\times 10^{-5}$

or 0.00006

$c = \lambda \nu$

$c = 3.00 \times 10^8$ m/s

$\nu = 5.00 \times 10^{12}$ Hz

5 __ 12

6×10^{-5}

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5 What type of radiation has a frequency of 8.6 x 10^11 s^-1 is

s^{-1}

8.6×10^{11}

B x-ray

C radiowave

A infrared

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6 What type of radiation has a wavelength of 4.2 nm?

A infrared

C radiowave

B x-ray

4.2×10^{-9} m

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7 What type of radiation has a frequency of 5.6 MHz?

A infrared
B x-ray

C radiowave

5.6×10^{6} Hz

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5.2 Quantum Theory and the Atom

I. Bohr's Model of the Atom

A. Dual-wave-particle model accounts for many scientific phenomena but leaves certain phenomena unexplained

B. Niels Bohr proposed a model of the atom that correctly predicted the frequencies of the lines of hydrogen's emission spectrum

1. In other words, Bohr determined that hydrogen atoms could only occur at certain energy levels

C. Ground state - the lowest allowable energy state of an atom

1. When an atom gains energy, it becomes excited

2. Bohr suggested electrons in atoms moves around in predictable circular orbits

D. Bohr assigned a number, n, known as the quantum number, to each orbital

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E. Hydrogen like Spectrum

1. When an atom of an element is in an excited state, the electron can drop from higher-energy orbits to lower-energy orbits

2. $\Delta E = E_{\text{higher energy orbit}} - E_{\text{lower energy orbit}}$

$\rightarrow E_{\text{photon}} = h\nu$

3. Because only certain atomic energies are possible, only certain frequencies of EM radiation can be emitted

F. As helpful as Bohr's model was, it did not explain EM spectrum of other elements

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II The Quantum Mechanical model of the Atom
A. Louis de Broglie (1892 - 1987)
1. proposed that particles of matter, including electrons, behave like waves
B. de Broglie equation - predicts that all moving particles have wave characteristics
1. wavelength $\lambda = \frac{h}{m \cdot v}$ — h planck's constant, m mass, v - velocity
C. Heisenberg uncertainty principle
1. Shows that it is impossible to take any measurement of an object without disturbing the object
D. Schrodinger wave equation
1. treated Hydrogen atom as a wave
2. Quantum mechanical model of the atom - atomic models where electrons are treated as waves
E. Electrons probable equation
1. Wave function predicts 3-D region around nucleus called atomic orbital, which predicts electron's probable location

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III Hydrogen Atomic Orbital
A. Principal quantum number - indicates the size and energy of atomic orbitals
1. As n increases, orbital becomes larger, electrons spend time farther away from the nucleus, the atom's energy increases
B. Principal energy level
1. Seven energy levels for H, n can = # 1-7
C. Energy sublevels
1. The number of energy sublevels in principal energy increase as $n \uparrow$
D. Shapes of orbitals
1. Sublevels are named for shape
a. s = spherical
b. p = dumbell-shaped
c. d, f — have differences in shapes

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5.2 Review

Grade: 11th

Subject: Chemistry

Date: 12/4

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1 According to the Bohr model of the atom, electrons move in circular orbits around the nucleus.

False

True

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2 How many energy sublevels are contained in Hydrogen's third level?

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3-D

3 Electrons occupy two-dimensional regions of space called atomic orbitals.

True

False

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4 The de Broglie equation relates a particle's wavelength to its______, its _______, and Planck's constant.

A mass, acceleration

C weight, acceleration

D weight, velocity

B mass, velocity

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5 The lowest allowable energy level of an atom is called its ground state.

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6 When an atom gains energy it is said to be in an excited state.

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5.3 Electron Configuration

I. Ground-state electron configuration-

A. Some atoms of heavy elements contain more than 100 electrons

B. The arrangement of electrons in an atom is called the atom's electron configuration

C. Aufbau principle- each electron occupies the lowest energy orbital available, Figure 18, p.300

D. Pauli exclusion principle- represents electrons orbitals w/ arrows in boxes

1. Spin of electron: $\boxed{\uparrow\downarrow}$ opposite directions

a. Max # of electrons related to each orbital (or principal energy level) equals $2n^2$

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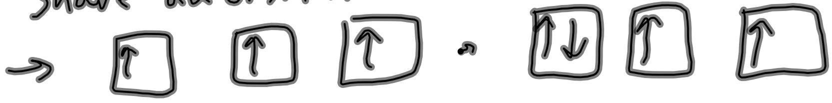
E. Hund's rule - since electrons have like charges, they fill equal energy orbitals in order before 2 electrons share an orbital



II Electron Arrangement

A. Atom's electron configuration can represented in two different ways; orbital digrams + electron configuration

B. Electron configuration notation -

1. designates the principal energy level and energy sublevel associated w/ each atom's orbitals

ex: $Na^{+} = 1s^{2}\,2s^{2}\,2p^{6}\,3s^{1}$

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☆ C. Noble-gas notation - puts last Noble gas in brackets [ ] and highest energy orbital

→ ex: [Ne] $1s^{1}$ = Na^{+}

D. Exceptions to predicted configurations:

1. For certain elements (Cr, Cu) are unique due to half-filled orbitals and filled sets of s and d orbitals

III Valence Electrons - electrons in outermost orbitals

A. Sulfur (S) has 6 valence electrons →

[Ne] $3s^{2}\,3p^{4}$

B. Electron-dot structures - consists of Elements symbol and dots surrounding it

see Table 6, p.308 ex: N = $\cdot\ddot{N}\!\!:$ → [He] $2s^{2}2p^{3}$

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5.3 Review

Grade: 11th

Subject: Chemistry

Date: 12/6

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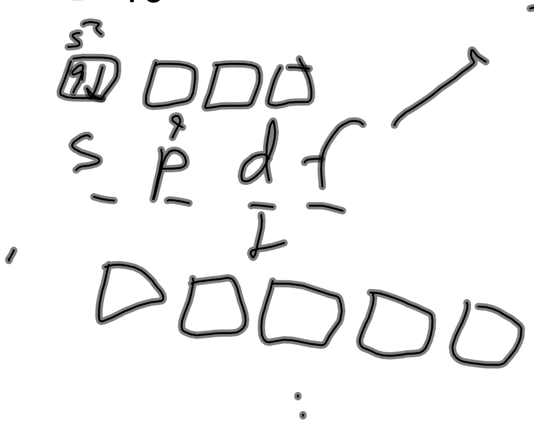
1 How many electrons can an orbital hold?

A 1

C 6

D 10

B 2



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2 The correct electron configuration for tin (Sn) is [Kr] $5s^2$ $3d^{10}$ $4f^{14}$ $5p^4$.

True

[Kr]$5s^2 4d^{10} 5p^2$

False

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3 What is an electron dot structure?

B an element symbol with a positive charge

C an element symbol surrounded by its innermost electrons

A An element surrounded by dots representing its valence electrons

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4 Why does the 4s orbital begin to fill before the 3d orbital begins to fill?

A s orbitals always fill before d orbitals

B the 4s orbital has higher energy than the 3d orbital

C s orbital always fill before d orbitals

D the 4s orbital has lower energy, therefore it fills before the 3d orbital

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5 Which of the following statements expresses Hund's rule?

A Electrons in orbitals must posses opposite spins

C electrons with the same spin fill all orbitals

D P orbitals may contain up to six electrons

B Single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbital

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